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ABSTRACT

This study guide is part of an interdisciplinary program of studies entitled the Science and Engineering Technician (SET) Curriculum. This curriculum integrates elements from the disciplines of chemistry, physics, mathematics, mechanical technology, and electronic technology with the objective of training technicians in the use of electronic instruments and their application. This guide provides that component of the content related to chemistry and provides an introduction to the following topics: (1) chemical laboratory safety and practice; (2) atomic structure; (3) inorganic chemistry; (4) nuclear chemistry; and (5) organic chemistry. (Author/SK)

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15. Abstract

This study is part of an interdisciplinary program of studies entitled the Science and Engineering Technician Curriculum (SET). This curriculum integrates elements from the disciplines of chemistry, physics, mathematics, mechanical technology, and electronic technology, with the objective of training technicians in the use of electronic instruments and their applications.

Two guides, Chemical Science and Technology I and II, provide that component of the content related to chemistry. This volume, provides an introduction to the following topics: (1) chemical laboratory safety and practices; (2) atomic structure; (3) inorganic chemistry; (4) nuclear chemistry; and (5) organic chemistry.

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CHEMICAL SCIENCE AND TECHNOLOGY I

A STUDY GUIDE OF THE SCIENCE AND ENGINEERING TECHNICIAN CURRICULUM

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Chapter I

Chemical Laboratory Safety and Practices

SECTION 1: ON THE PREVENTION OF EXPLOSIONS, FIRES AND GREAT BODILY HARM

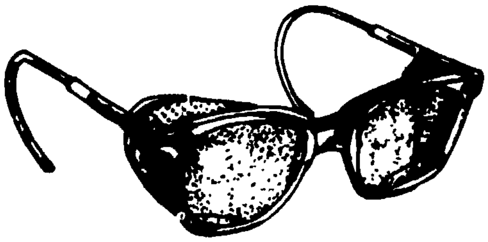
Common to all physical activities is the need to work safely. Since this is possibly your first exposure to a chemical laboratory environment, a few basic "Nevers" should be presented. The hazards of laboratory work are reduced in proportion to your understanding of safe practices.

1. Never work in a chemistry laboratory without proper eye protection. Contact lenses should never be worn in a chemical laboratory.

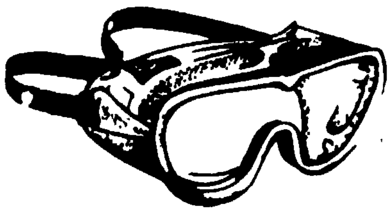
2. Never work alone in a laboratory.

3. Never perform unauthorized experiments in the laboratory or at home.

4. Never use chemicals without carefully reading the labels - many accidents have been caused because certain chemicals have similar names. Sodium chloride (table salt) is absolutely necessary for man's existence while elemental sodium will explode in water and chlorine was used in World War I as a trench gas.



Safety Glasses



Safety Goggles

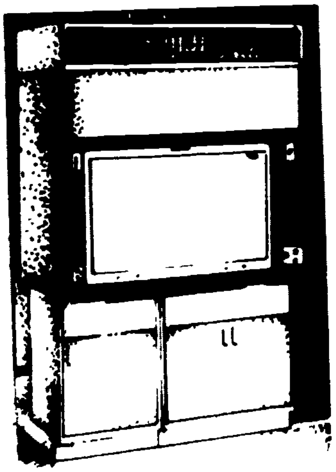
5. Never return excess chemicals to their original container. They should be discarded (usually down the drain with an excess of water or into a special waste chemicals container). Consult your instructor or supervisor when in doubt. The problem of disposing of excess chemicals can be minimized by taking only the amount of material required for the experiment.

6. Never perform any experiments or reactions that produce objectionable or unknown gases without using a fume hood.

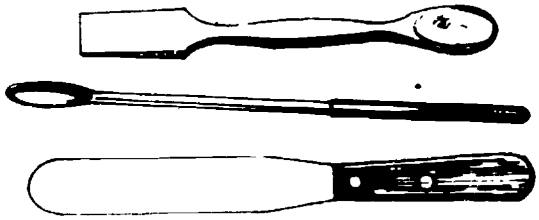
7. Never use any chemicals found in an unlabeled container and conversely never store chemicals in a container without labeling it.

8. Never bring chemicals into contact with the skin - use a spatula or other sampling device.

Safety Container



Fume Hood

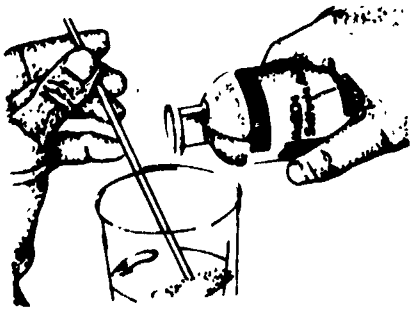


Spatulas

9. Never use flammable chemicals near an open flame. Smoking is not permitted in most chemical laboratories.

10. Never pour water into concentrated acid as it might spatter and/or break the container because of the excessive heat generated.

"While we've got a good view, I'd like to point out some other places where you shouldn't strike matches."

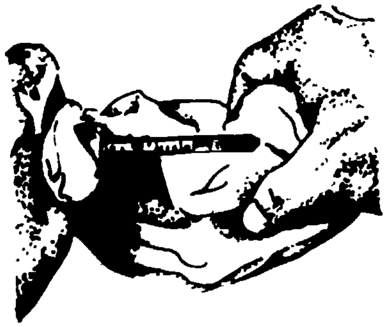


Always pour **acid** into water very slowly and stir constantly.

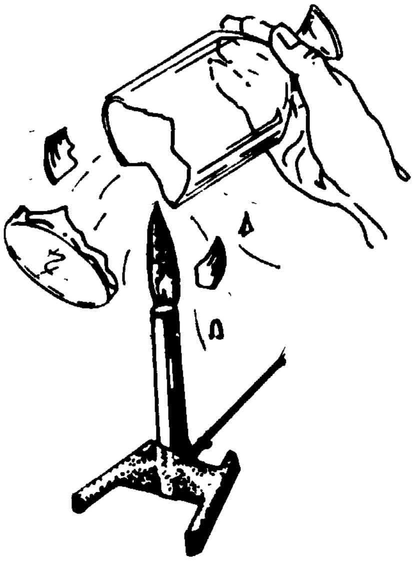
11. Never attempt to insert a glass tube or thermometer into a stopper without a lubricant (water or glycerine). The glass tube should be held in a cloth to minimize the hazard if the tube does break.

12. Never fail to clean up your laboratory area, replace chemicals, and turn off all gases.

13. Never heat "soft" glass containers (bottles, funnels, thick-walled glassware, etc.) in an open flame. This type of glass is not manufactured to withstand high temperatures or thermal shock.



Always lubricate a thermometer



Not all glassware can be heated.

14. Never work in a laboratory without first learning the operation and/or location of all fire extinguishers, fire blankets, safety showers, eye wash fountains, and exits.

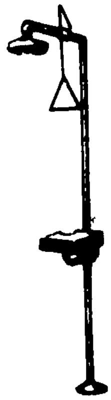
15. Never point the open end of a test tube at yourself or any other person while the tube is being heated or during a reaction.

16. Never taste a laboratory chemical. If instructed to smell a chemical, do so by gently fanning the vapors toward your nose.

17. Never use your mouth to pipet chemicals; always use a rubber bulb.

18. Never use compressed air for jokes or point the hose at another person. Practical jokes using compressed air have resulted in severe injury and sometimes death.

"By the way, Ed, that's a perchloric acid mixture in that Coke bottle."



Safety Shower

Occupational Safety and Health Act of 1970

The Williams - Steiger Occupational Safety and Health Act was approved by Congress and became effective April 28, 1971. It is now generally known as OSHA.

Under the law "each employer (must) furnish his employees a place of employment free from recognized hazards that might cause serious injury or death;" and the Act further requires that employers comply with the specific safety and health standards issued by the Department of Labor. The employee has the duty to "comply with the specific safety and health standards, rules, regulations, and orders issued under the act and applicable to his conduct."

Toxicity

The study of the properties and effects that chemicals have on living systems is pharmacology. Toxicology is the branch of pharmacology that deals with poisons.

Toxicity is evaluated from studies performed on animals and extrapolated for application to the human body. This procedure is sometimes lacking in medical validity but is absolutely necessary until a better procedure becomes available. The three basic units of toxicity are:

1. lethal dose, 50 percent kill, abbreviated LD_{50}. The amount of a material which, when administered to laboratory animals, such as white mice or guinea pigs, kills half of them. It is

expressed in units of mg/kg, the milligrams of material administered per weight, in kilograms, of the animal. Assuming that a substance is just as lethal to humans as it is to white mice, the lethal dose for an average person who weighs w kilograms is merely LD_{50} X w.

2. lethal concentration, 50 percent kill, abbreviated LC_{50}. The concentration of a material, normally expressed as ppm by volume which, when administered to laboratory animals, kills half of them in some time period of exposure. Parts per million by volume is equivalent to ppm in the number of molecules present, since like numbers of molecules of different gases occupy almost the same space.

3. threshold limit value, abbreviated TLV, is the upper limit of a toxicant concentration to which an average healthy person may be repeatedly exposed on an all-day, everyday basis without suffering adverse effects. It is usually expressed as ppm for gases in air or alternatively milligrams per cubic meter (mg/m^3) or micrograms per cubic meter ($\mu g/m^3$) for fumes and mists in air. The TLV was formerly known as the maximum allowable concentration. Values of the TLV are set by the American Conference of Governmental Industrial Hygienists (ACGIH) and are periodically revised when new information necessitates a change. The respiratory system is the most likely route for a toxic substance to accidently enter the body. An asphyxiant would be a substance that can arrest the respiratory system to some degree. Smoke presents a serious hazard in that the finely divided particles of carbon (soot) obscure vision, coat bronchial surfaces, actually absorb toxic gases, and contain complex hydrocarbons which could be carcinogens (cancer-producing substances). See the table in this chapter for toxicity data on some common chemicals.

Laboratory

a. The student should be able to restate each "never" in his own words and briefly describe the hazards involved.

b. The procedures for disposing of various types of excess chemicals should be demonstrated and explained by the instructor.

c. The proper utilization of a fume hood should be shown.

d. The use of various types of spatulas and other sampling devices for solids should be practiced.

e. The sizing and/or boring of a stopper with the subsequent insertion of a thermometer into the (lubricated) stopper should be practiced.

f. The proper utilization of available fire blankets, safety showers and eye wash fountains should be demonstrated.

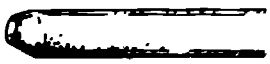
Student Problems

1. What will happen if water is poured directly into concentrated sulfuric acid?
2. An experiment that generates chlorine gas should be carried out under what conditions?
3. What does OSHA mean?
4. In your own words define LD_{50}.

SECTION 2: CHEMISTRY LABORATORY GLASSWARE, HARDWARE, AND BEWARE

Glassworking

In addition to learning the laboratory safety rules, it is necessary to become familiar with the equipment that a laboratory worker might be expected to use that is commonly used in most chemistry labs. Since many chemicals are quite reactive, most laboratory equipment is composed of inert substances like platinum, polyethylene, porcelain and glass. Many laboratory procedures call for some relatively simple glass tubing construction. All glass tubing for example, must be "firepolished" and lubricated (H_2O or glycerine) before insertion through a stopper. Firepolishing is accomplished by simply rotating the tubing end in an open flame until the sharp edge is rounded.

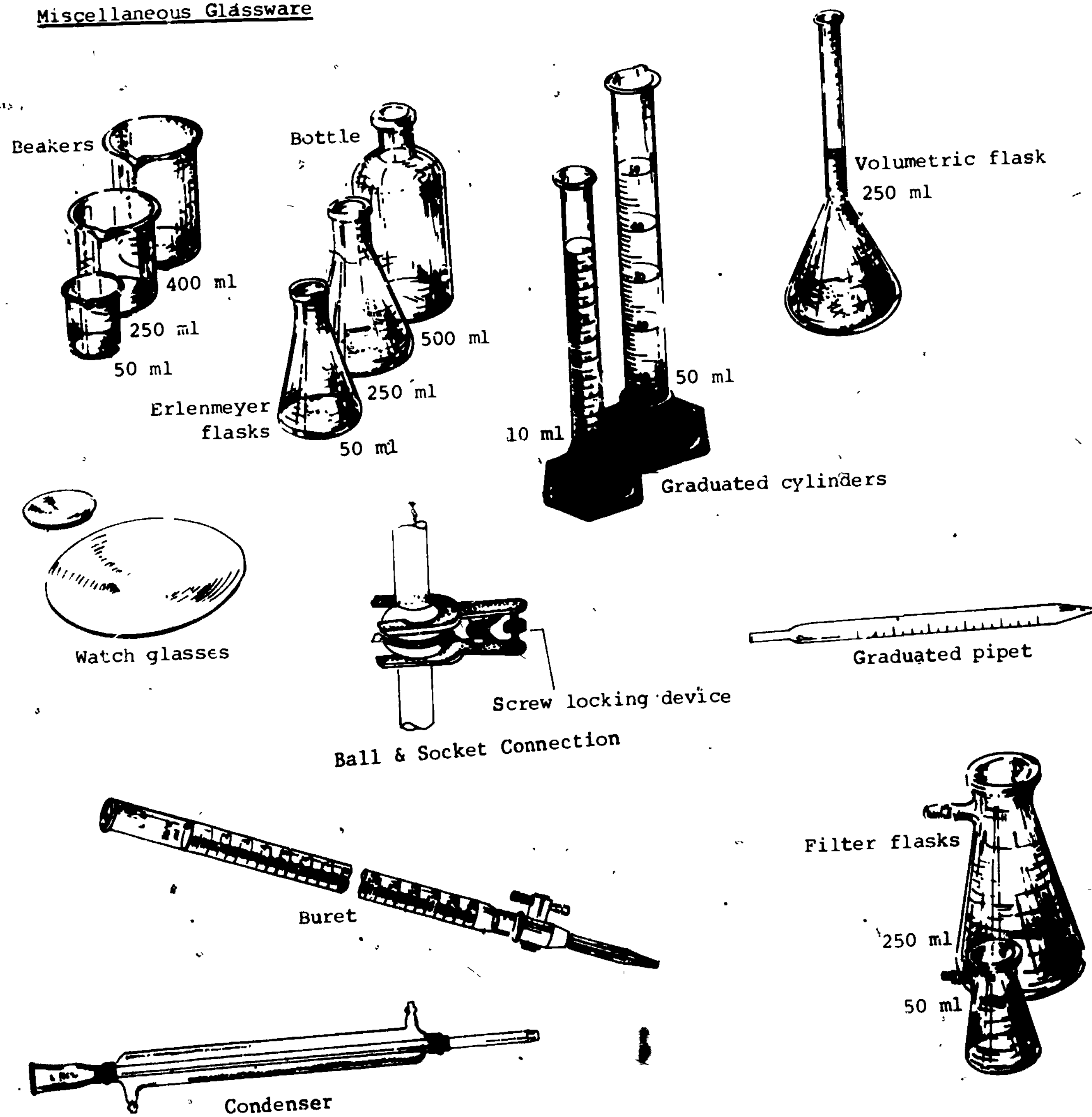


Firepolished too long

Correct Firepolish

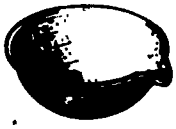
There are many other glass working techniques that are required of a competent laboratory worker. Expertise should be gained in the construction of U-tubes, bends, flares, ball and socket connections, standard tapers, metal to glass joints, glassware cleaning, etc.

Miscellaneous Glassware



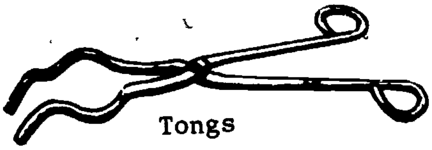
Miscellaneous Hardware

Evaporating dish

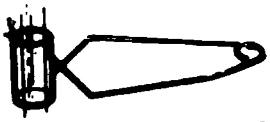


Crucible and lid

Mortar and pestle

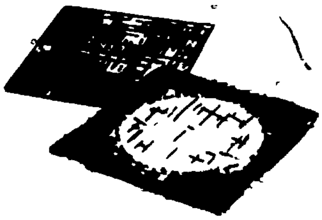


Tongs



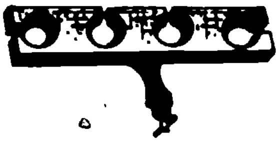
Test tube holder

Buchner funnel



Wire gauze with
asbestos center

Flame spreader



Funnel support

Laboratory

The student should be able to cut and "firepolish" a glass tube. He should be able to identify the various pieces of laboratory glassware and hardware described in this section. If time permits, some additional practice should be gained in constructing U-tubes and bends. A demonstration of ball and socket connections and standard taper glassware would be helpful.

Student Problems

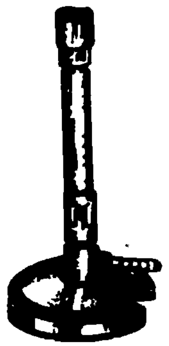
1. Name the four most common substances used in the construction of chemistry laboratory equipment.

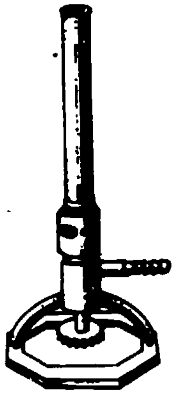
2. Why is it necessary to firepolish a glass surface?

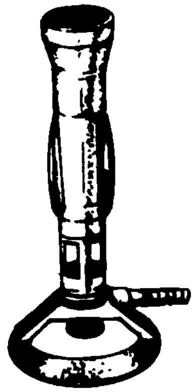
3. State the purpose of a watch glass.

SECTION 3: BURNERS, OVENS, AND OTHER HOT THINGS

The most common heating device in a chemistry laboratory is the Bunsen burner. A simple burner mixes natural gas and air in the barrel and is ignited at the top. The air is drawn into the barrel through an adjustable opening near the base. This type of burner can reach temperatures near 850^{o} C. A Tirrill burner is simply a Bunsen burner with an adjustable gas orifice. The Meker (Fisher) is constructed with a grid on the top and can reach much higher temperatures.

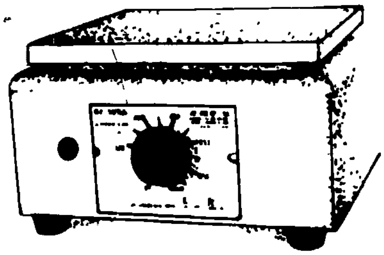

Bunsen

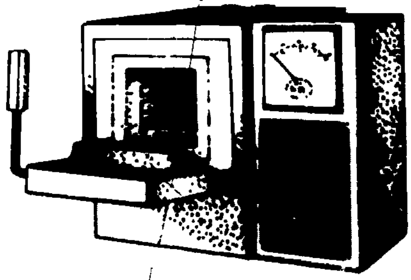

Tirrill


Meker

Many chemicals (especially organics) are too flammable to allow heating with an open flame as found in conventional burners. A hot plate, heating mantle, or steam bath should be used if a fire hazard exists. Conventional oven, vacuum ovens, and muffle furnaces are also available for heating chemicals to various temperatures. Even an oven might not be safe for heating volatile and flammable substances because of the confinement of vapors and the exposure of electrical connections.


Heating mantle


Hot Plate


Muffle Furnace

Flammable Chemicals

The minimum concentration (normally expressed in per cent by volume) of gas or vapor in air below which a substance does not burn when exposed to an ignition source is called the lower explosive limit (LEL%). The maximum concentration of the substance in air above which ignition does not occur is called the upper explosive limit (UEL%). For example, the lower explosive limit of acetone is 2.6 percent by volume in air, whereas the upper explosive limit is 12.8 percent in air. A mixture of acetone vapor and air having a concentration of less than 2.6 percent in air is too "lean" in fuel vapor to burn. Similarly, a mixture containing more than 12.8 percent acetone vapor in air is too rich in fuel to burn.

Liquids must be heated to a temperature where the production of vapor yields a concentration within its flammable range before it can ignite. There is a minimum liquid temperature at which a spark or flame cause an instantaneous flash in the vapor space above the liquid. This temperature is called the flash point. Flash point can be defined as the minimum temperature of the liquid at which it gives off vapor sufficient to form an ignitable mixture with the air near the surface of the liquid within the vessel used. Flash points are determined by several methods by the American Society for Testing and Materials (ASTM): the Tagliabue and Pensky-Martens Closed Cup methods and the Cleveland Open Cup method. In each of these methods a small pilot flame is passed over a sample of the liquid contained in an open or closed vessel at a specified temperature. The flash point is the temperature at which ignition of the liquid first occurs. In general, the Open Cup method gives results 3° to 6°C higher than the Closed Cup methods. The results obtained by the Open Cup method more nearly simulate the combustion of liquids under fire conditions.

As the temperature of a liquid is increased beyond the flash point, a temperature is reached at which the liquid gives off enough vapor to continue to burn when ignited. This temperature is called the fire point. Generally, the fire point is 15° to 30°C higher than the flash point. Finally, if the temperature of a liquid is increased beyond the fire point, a temperature is reached at which ignition occurs without the introduction of an ignition source. This temperature is called the autoignition temperature.

The following table contains some typical flammability properties of common solvents.

TOXICITY AND FLAMMABILITY OF SOME COMMON SOLVENTS

Common Solvent Names	Boiling Point (°F)	Flash Point (°F)	Explosive Limits		Auto-ignition Temp. (°F)	Threshold Limit Value		Major Health Hazard
			LEL%	UEL%		ppm in Air	mg/m^3 in Air	
Acetone	134	0	2.6	12.8	1000	1,000	2,400	Skin irritant: Headache
Benzene (benzol)	176	12	1.4	8.0	1000	25 (skin)	80 (skin)	Toxic by inhalation and skin absorption
n-butyl acetate	259	92	1.7	15.0	740	150	710	Irritating to eyes and respiratory tract; narcotic
butyl alcohol	244	114	1.7	18.0	650	100	300	Irritating to eyes, nose, and throat; causes headache and dizziness
Carbon tetrachloride	170	None	-	-	-	10 (skin)	65 (skin)	Narcotic; can cause organ damage
Chloroform	142	None	-	-	-	25	120	Anesthetic; causes eye irritation
Cyclohexanone	312	147	1.1	-	847	50	200	Eye and throat irritation, mild narcotic
Diethyl ether	94	-49	1.85	48	356	-	-	Anesthetic, causes eye irritation
Ethyl acetate	171	40	2.7	11.5	900	400	1,400	Irritating to eyes and respiratory passages; mild narcotic
Ethyl aclohol (ethanol)	173	61	3.3	19.0	750	1,000	1,900	Irritant narcotic
Ethylene dichloride	182	70	6.2	15.9	840	50	200	Strong irritant
Ethylene glycol monobutyl ether	340	141	-	-	472	50	240	Irritant
Heptane	209	25	1.2	6.7	452	500	2,000	Irritating to respiratory tract, mild narcotic
Hexane	156	<-10	1.2	6.9	500	500	1,800	Irritant
Isobutyl alcohol	226	100	1.7	-	800	100	300	Strong irritant
Isopropyl alcohol	180	70	2.5	5.2	852	400	980	Irritant, mild narcotic
Kerosene	347-617	100-165	1.16	6.0	490	-	-	Irritant, headaches, moderate narcotic effect
Methyl acetate	136	14	4.1	13.0	935	200	610	Narcotic and irritant
Methyl alcohol (methanol)	149	65	6.0	36.5	878	200	260	Strong narcotic and mild irritant
Methylene chloride	104	None	15.5 in O_2	66.4 in O_2	1224	500	1,740	Dangerous to eyes, narcotic
Methyl ethyl ketone (butanone)	175	22	1.8	11.5	960	200	590	Irritant and narcotic
Methyl isobutyl ketone (hexone)	244	73	1.3 at 122°F	8.0 at 212°F	858	100	410	Irritating to eyes and mucous membranes
n-propyl alcohol	207	59	2.5	13.5	812	200	500	Irritant
Propylene oxide	93	-35	2.1	21.5	-	100	240	Moderate irritant
Tetrahydro-furan (THF)	150	1	2.3	11.3	-	200	590	Irritating to eyes and mucous membranes, moderate narcotic
Toluene (toluol)	231	40	1.27	7.0	1026	100	370	Moderate narcotic, can cause organ damage
Trichloro-ethane-1,1,1	165	None	-	-	-	350	1,900	Moderate narcotic and irritant
Trichloro-ethylene	189	None	-	-	770	100	535	Narcotic and addictive
Turpentine	309-338	95	0.8	-	488	100	560	Mild cause of allergy, toxic
VM & P naphtha	212-284	20	0.9	6.0	450	100	400	Causes intoxication
Xylene (Xylol)	280	77	1.1	7.0	924	100	435	Moderate irritant

Fire Extinguishers

There are three principal ways to extinguish a fire: 1) remove the fuel, 2) remove the oxidant, and 3) lower the temperature.

Underwriters Laboratories has developed a classification system which has been adopted by the National Fire Protection Association. The classifications of fire are as follows:

Class A - Fires in ordinary combustible materials such as wood, paper, coal or fabrics.

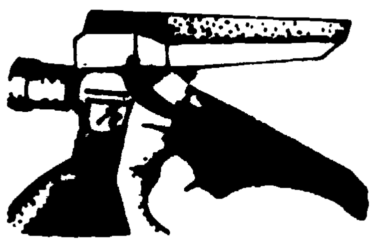
Class B - Fires in flammable liquids and gases.

Class C - Fires in or near electrical apparatus, where a non-conducting extinguishing agent is needed.

Class D - Fires in combustible metals, such as magnesium, potassium, sodium, lithium and others.

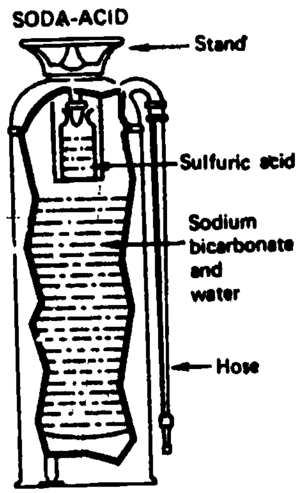
Carbon Dioxide Extinguishers (B, C)

This extinguisher consists of a high pressure cylinder containing liquid carbon dioxide under a pressure. Carbon dioxide extenguishers are recommended for use on Class B and C fires. To operate this extinguisher, pull the pin, direct the horn at the base of the fire, squeeze the handle and sweep the horn back and forth covering the full width of the fire. The effective range is 3 to 8 feet and the discharge time 20 to 30 seconds.



"Pulling the pin" on a Fire Extinguisher

All-purpose Dry Powder Fire Extinguisher



Soda-Acid Fire Extinguisher

Carbon Dioxide Fire Extinguisher

Soda-Acid Extinguishers (A, B)

The common soda-acid fire extinguisher contains sodium bicarbonate and sulfuric acid solutions. When inverted, the acid and bicarbonate mix and react violently, forming carbon dioxide, which forces the solution out of the nozzle with great pressure. This type of fire extinguisher is not utilized as much now that the all-purpose dry powder types are available.

All-Purpose Dry Powder (A, B, C)

This extinguisher consists of a substantial tank containing dry powder under nitrogen pressure. The powder consists mainly of monoammonium phosphate. This extinguisher is effective on Class A, B, and C fires. To operate this extinguisher pull back on the ball release on the handle, direct the nozzle at the base of the fire and squeeze the handle. Discharge time is 20 to 30 seconds with an effective range of 10 to 20 feet.

Met-L-X Dry Powder (B, C, D)

The Met-L-X dry-powder extinguisher is identical in appearance to the ordinary dry-chemical extinguisher except that it is yellow in color versus red. Operation is identical. The powder consists primarily of sodium chloride. Discharge time is 20 to 30 seconds with an effective range of 5 to 10 feet. It is effective on Class B and C fires but is designed primarily for Class D fires. When using this extinguisher, the powder should be permitted to gently fall on the burning metal surface. Directing the powder stream directly at the burning metal will only scatter the fire.

Pressurized Water Extinguishers (A)

This extinguisher simply consists of a tank containing water under pressure. Pressurized water extinguishers are to be used on Class A fires only. Under no circumstances should they be used on Class B, C, or D fires since the water could spread or increase the intensity of the flames. Class C fires could present an electrical hazard with water.

The effective range is 20 to 30 feet and a discharge time of 45 to 60 seconds.

To operate this extinguisher, pull the pin, direct the nozzle at the fire and squeeze the handle.

Laboratory

The student should be able to match fire extinguishers to the various types of fire. He should be able to connect, ignite, and adjust all three burner types (Bunsen, Tirrill, and Meker). He should have a working knowledge of a heating mantle and muffle furnace.

Student Problems

1. What is the purpose of a heating mantle?

2. Define flash point.

SECTION 4: PRESSURE, PRESSURE EVERYWHERE

Many laboratories use compressed liquids and gases which are contained under very high pressures (typically 100 to 2500 psi) in metal cylinders. These substances are potentially very dangerous because they are pressurized, flammable, corrosive, toxic, and/or extremely cold. Many industrial accidents have occurred from the mishandling of these cylinders and their contents. Always consult the manufacturer's data sheet on handling procedures or special properties of these compressed gases and liquids.

All gases can be reduced eventually to liquids and the liquids to solids by an appropriate decrease in the applied temperature and/or an increase in the applied pressure. However, there is a temperature for all gases above which pressure alone cannot condense them to liquids. This temperature is known as the critical temperature. For example, carbon dioxide easily liquefies if a sufficient pressure is applied while the gas is below 88°F. No applied pressure can cause liquefication of the gas above this temperature. The pressure required to liquefy a gas at its critical temperature is called the critical pressure. The critical pressure of carbon dioxide is 73 atm; that is, the application of 73 atm of pressure causes carbon dioxide at 88°F to liquefy. The volume occupied by a material at its critical pressure and critical temperature is called its critical volume. The critical volume of one mole (44 g) of carbon dioxide is 0.095 l. Critical data for several other gases are given below.

CRITICAL TEMPERATURES AND PRESSURES

Substance	Critical Temperature °C	Critical Temperature °F	Critical Pressure atm	Critical Pressure psi	Storage Temperature
Ammonia	130	266	115	1,691	-33°C
Carbon dioxide	31.1	88	73	1,073	-57°C
Diethyl ether	197	387	35.8	526	--
Hydrogen	-234.5	-390	20	294	-254°C
Nitrogen	-146	-230.8	33	485	-196°C
Oxygen	-118	-180	50	735	-183°C
Sulfur dioxide	155.4	311.7	78.9	1,160	-10°C

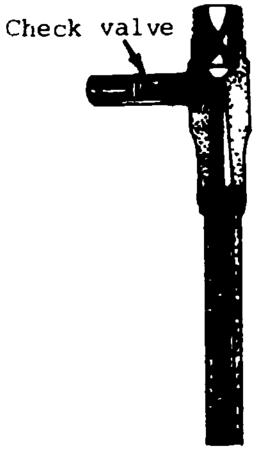
Gases that are liquidified by applying pressure and stored in insulated containers are referred to as cryogenic liquids. These very cold liquids require specially designed vessels for storage. A common device is the Dewar flask which is a double-walled glass vessel with a silvered interior, similar to the lining of a "Thermos" jug. The extreme cold temperatures and relatively high pressures of these cryogens present many hazards:

a. Damage to living tissue. Contact with cryogenic fluids can cause a localized solidification of tissue and produce a burn as painful as that received from a heat source. The extreme coldness causes a local arrest in the circulation of the blood, and any skin tissue that has been exposed to cryogenic fluids should be restored to normal body temperature as rapidly as possible. This is most easily accomplished by merely immersing the damaged skin area into water around 108°F.

b. A high expansion rate on vaporization. Liquid methane, for example, expands to approximately 630 times its initial volume when it vaporizes. If the mechanism of cooling methane fails or is inadequate to keep it fluid, the internal pressure within the storage container increases so tremendously that rupture of the container is likely unless proper venting has been provided.

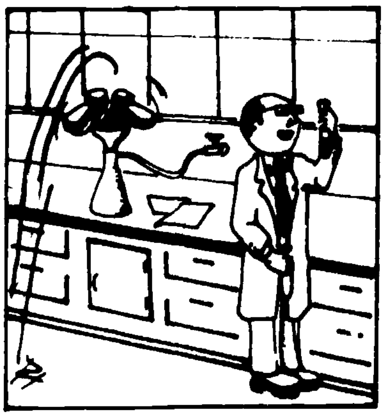
c. An ability to liquefy other gases. Recognizing that cryogenic fluids are so extremely cold, it is logical to expect that they are also capable of condensing and even solidifying other gases. Gaseous air, for example, solidifies when it is exposed to a number of cryogenic fluids. The solidification of air is a major hazard if it occurs in venting tubes of the storage container, since it blocks the passage and prevents the release of pressure buildup.

A low pressure (20 torr range) can be created in a laboratory by using a water aspirator (Bernoulli's Principle). A high vacuum source can be found in a mechanical vacuum pump. Vacuum systems should be treated with the same respect as pressurized systems; pressurized systems can explode while high vacuum systems can implode with the same consequences. A few "Never" rules on the handling of compressed gases, liquids, and vacuum systems should serve as a guide or listing of "behavioral objectives" that you should develop expertise in before leaving this section.



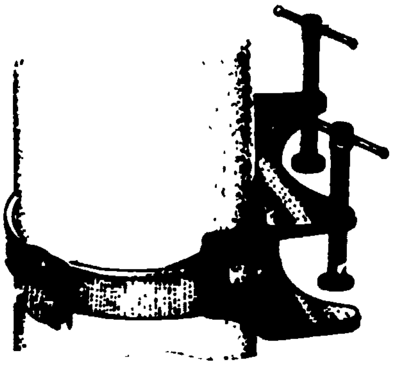
Water Aspirator

Mechanical Vacuum Pump

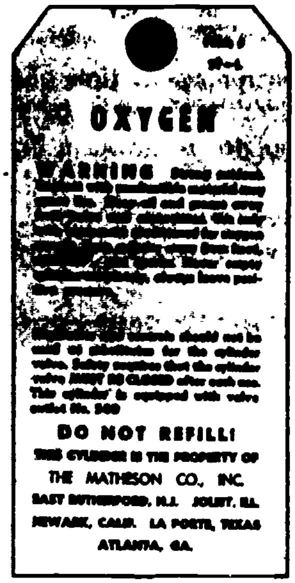


"How's the suction on that filter now?"

1. Never use or work with a cylinder that is not securely fastened or chained to a stationary support.



Gas Cylinder Clamp



Warning Tag

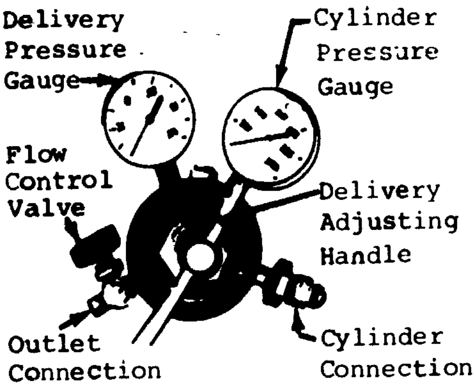
2. Never "adapt" an improper regulator to a cylinder since this bypasses a built-in safeguard of matching cylinders with only specific regulators.

3. Never use a regulator on an oxygen cylinder that has been used on some other gas. The high pressure oxygen can react explosively with any residual oil.

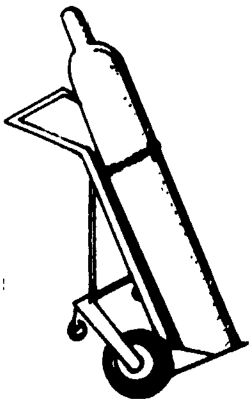
4. Never stand directly in front of a regulator when operating the high pressure valve.

5. Never transport a cylinder by simply rolling. It should be securely fastened in a special cylinder transport cart.

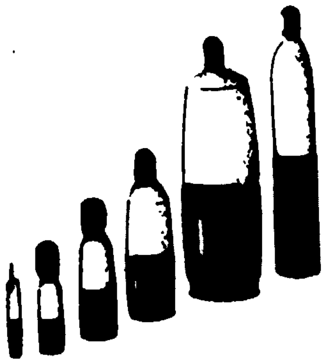
6. Never leave a high pressure, vacuum, and/or corrosive gas system on when not in use.



Gas Cylinder Regulator



Cart for Transporting Gas Cylinders



Cylinders of Compressed Gases

7. Never store gas cylinders in bright sunlight or near heat sources. Cylinders which contain hydrocarbons or other highly flammable substances should be stored out-of-doors.

8. Never use a Dewar flask without wrapping the vessel in tape or insulation to minimize any implosion hazard.

Laboratory

The instructor should demonstrate the proper handling techniques involved with cryogenic liquids. The student should set-up and operate both a water aspirator and a mechanical vacuum pump. These set-ups should include the installation of safety bottles to prevent back-up. The proper handling of compressed gas cylinders should be first demonstrated, then practiced by each student (to include: transporting, securing, attachment of proper regulator, and adjustment of the regulator).

Student Problems

1. What is a cryogenic liquid?

2. A conventional water aspirator is capable of pulling ________________ torr under optimum condition.

3. What special problem does compressed oxygen present?

SECTION 5: THE LABORATORY NOTEBOOK VERSUS THE PAPER TOWEL

The success or failure of an experiment can depend on the thorougness and accuracy of the records. How detailed should a laboratory notebook be? A simple rule would be to include everything that might be necessary for another person to duplicate your experiment directly from the accounts in your notebook. A cardinal rule is to enter data and observations exactly as you observed them not what you think they should be. Most companies have specific procedures concerning technicians' notebooks, since these volumes contain company records and establish patent rights.

A preliminary notebook guide is given in the following outline. The list is by no means complete since both instructors and companies vary on exact format and entries required.

1. Use only permanently bound books, no loose leaf notebooks.

2. Never erase; corrections are made by drawing a single line through the erroneous material and entering the correction in the margin with your initials and date.

3. All notebook entries should be made in ink and witnessed by at least one additional laboratory worker.

4. The notebook should contain a "Table of Contents" and at least the following information:

   a) All pages numbered consecutively and all participants identified.

   b) Every experiment titled and dated.

   c) Brief objective statements for each experiment.

   d) Equations (chemical, mathematical, etc.) should be included.

   e) Make cross references to related materials in the same or other notebooks.

   f) Tables to include all principle reactants and products, their formulas, and important physical properties.

   g) The exact procedure followed.

   h) All data and observations.

   i) Calculations, especially set-ups and equations used.

   j) Conclusions and/or interpretations.

Laboratory

The student should be expected to maintain a laboratory notebook throughout this course and to submit it for periodical evaluation.

Student Problems

1. Why is a notebook so important in industry?

Chapter II

Atomic Structure

SECTION 1: ATOMS, ELEMENTS AND ATOMIC WEIGHTS

All matter in our universe is composed of one or a combination of more than one element (listed below). An atom is the smallest unit of an element. Each pure element is composed of a unique type of atom, which is characteristic of that element, and contains no other type of atom. A compound can be defined as matter made of two or more elements in which the elements always occur together in a definite proportion by weight. (H_2O, NaCl, H_2SO_4, etc.). A mixture is made of two or more elements or compounds, but not in definite or constant proportion. Soil is a good example of a mixture.

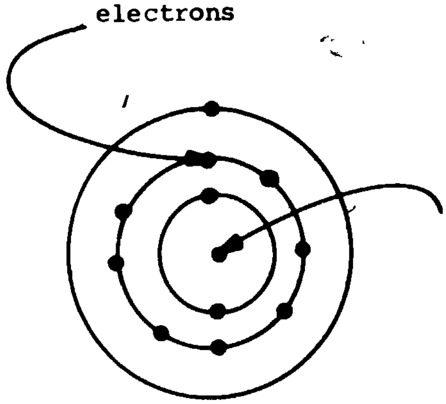
ELEMENT	SYMBOL	ATOMIC NUMBER
Actinium	Ac	89
Aluminum	Al	13
Americium	Am	95
Antimony	Sb	51
Argon	Ar	18
Arsenic	As	33
Astatine	At	85
Barium	Ba	56
Berkelium	Bk	97
Beryllium	Be	4
Bismuth	Bi	83
Boron	B	5
Bromine	Br	35
Cadmium	Cd	48
Calcium	Ca	20
Californium	Cf	98
Carbon	C	6
Cerium	Ce	58
Cesium	Cs	55
Chlorine	Cl	17
Chromium	Cr	24
Cobalt	Co	27
Copper	Cu	29
Curium	Cm	96
Dysprosium	Dy	66
Einsteinium	Es	99
Erbium	Er	68
Europium	Eu	63
Fermium	Fm	100
Fluorine	F	9
Francium	Fr	87
Gadolinium	Gd	64
Gallium	Ga	31
Germanium	Ge	32

Element	Symbol	Atomic Number
Gold	Au	79
Hafnium	Hf	72
Helium	He	2
Holmium	Ho	67
Hydrogen	H	1
Indium	In	49
Iodine	I	53
Iridium	Ir	77
Iron	Fe	26
Krypton	Kr	36
Lanthanum	La	57
Lawrencium	Lw	103
Lead	Pb	82
Lithium	Li	3
Lutetium	Lu	71
Magnesium	Mg	12
Manganese	Mn	25
Mendelevium	Md	101
Mercury	Hg	80
Molybdenum	Mo	42
Neodymium	Nd	60
Neon	Ne	10
Neptunium	Np	93
Nickel	Ni	28
Niobium	Nb	41
Nitrogen	N	7
Nobelium	No	102
Osmium	Os	76
Oxygen	O	8
Palladium	Pd	46
Phosphorus	P	15
Platinum	Pt	78
Plutonium	Pu	94
Polonium	Po	84
Potassium	K	19

ELEMENT	SYMBOL	ATOMIC NUMBER
Praseodymium	Pr	59
Promethium	Pm	61
Protactinium	Pa	91
Radium	Ra	88
Radon	Rn	86
Rhenium	Re	75
Rhodium	Rh	45
Rubidium	Rb	37
Ruthenium	Ru	44
Samarium	Sm	62
Scandium	Sc	21
Selenium	Se	34
Silicon	Si	14
Silver	Ag	47
Sodium	Na	11
Strontium	Sr	38
Sulfur	S	16
Tantalum	Ta	73
Technetium	Tc	43
Tellurium	Te	52
Terbium	Tb	65
Thallium	Tl	81
Thorium	Th	90
Thulium	Tm	69
Tin	Sn	50
Titanium	Ti	22
Tungsten	W	74
Uranium	U	92
Vanadium	V	23
Xenon	Xe	54
Ytterbium	Yb	70
Yttrium	Y	39
Zinc	Zn	30
Zirconium	Zr	40

Alphabetical Listing of the Elements

A visual picture of an atom is based on the assumptions of Niels Bohr (early 1900s) that electrons in an atom are confined to definite energy states or shells, and that if an electron moves from one level to a higher one, the atom has absorbed an amount of energy exactly equal to the difference of the two states. If an electron drops to a lower shell, it will emit the energy (sometimes as visible light) corresponding to the difference. Bohr's model is the solar system model: the nucleus is like the sun; the electrons are like the planets and they move in definite orbits. Chemists no longer believe in the existence of fixed, ellipsoidal orbits; but they do believe in the existence of distinct energy states. Instead of corresponding to ellipsoidal orbits, the energy states are better described as peculiarly shaped regions called orbitals clustered near the atom's nucleus.



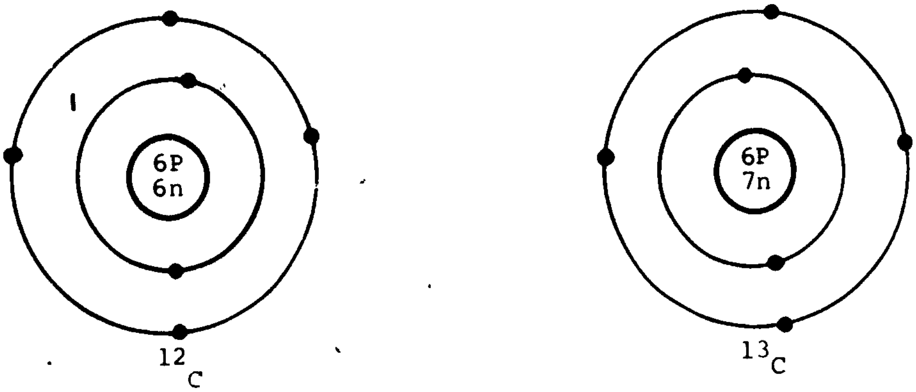
Bohr Model
of the atom

Atoms are composed of a relatively massive nucleus containing positively charged particles (protons) and neutral particles (neutrons) with negatively charged particles (electrons) of very small mass outside the nucleus. A neutral atom will always have as many protons as electrons.

Subatomic Particles			
	Symbol	Relative Mass	Charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	0*	-1

*Actually 0.05% of the mass of proton

The <u>octet rule</u>, says that atoms interact to change the number of electrons in their outer electronic shells in an attempt to get an electronic structure similar to that of a noble gas. The noble-gas structure consists of eight electrons in the outermost shell for all elements except hydrogen and helium, where a complete shell consists of only two electrons. The noble gases all have full quotas of electrons in their shells and this structure accounted for the chemical inertness of these elements. It is the tendency to achieve electronic structures similar to the noble gases that explains chemical bonding in all compounds. There are many exceptions to this octet rule, but the exceptions also can be explained by using the general principles of electron interactions.



Isotopes of Carbon

An isotope may be specified by writing the symbol for the element with the mass number (total number of protons and neutrons combined) as a superscript to the left. For example, the two principal isotopes of carbon that occur in nature have masses of 12.000 a.m.u. and 13.0033 amu. These two isotopes would be designated as ^{12}C and ^{13}C, respectively. A formal definition of isotope is given on page 27.

An atomic mass unit (amu) is defined as exactly 1/12 of the mass of the ^{12}C isotope. This means that the ^{12}C isotope is the standard against which all other atomic weights are determined; it has been assigned a mass of exactly 12.0000····. The average sulfur (S) atom is 2.672 times as massive as an atom of ^{12}C. The atomic mass of sulfur would be calculated as follows:

mass of ^{12}C x mass ratio of element = atomic mass of element

12.0000 amu x 2.672 = 32.06 amu S

Average Atomic Weight

The atomic weight of an element is the mass of an average atom of the element occurring in nature. The mathematical formula is:

$$\text{atomic weight} = f_1A_1 + f_2A_2 + \ldots$$

where f_1, f_2... are the fractional abundances of isotopes having the masses A_1, A_2,...

This convention is used because in everyday life we never deal with one atom at a time, but with vast numbers of them, and in doing ordinary calculations it is the average mass that is important.

Worked Example

In a natural sample of carbon, 0.989 of it is composed of ^{12}C (12.000 amu) and 0.0110 of is ^{13}C (13.0033 amu). The average atomic weight of carbon would be calculated as follows:

$$\text{Average atomic weight} = f_{12}A_{12} + f_{13}A_{13}$$

$$= (0.989)(12.000) + (0.011)(13.0033)$$

$$= 12.011$$

rounded off to the number of significant figures warranted by the data as given = 12.0.

Laboratory

The student, using devices like melting point capillaries, pycnometers, hydrometers, etc., should become familiar with measuring physical properties of elements, compounds, and mixtures.

Student Problems

1. How many neutrons would be found in an Argon (Ar) atom with an atomic mass of 40?

2. Naturally occurring potassium consists of 93.10% ^{39}K (38.98 amu), 6.88% ^{41}K (40.97 amu), and 0.01% ^{40}K (39.98 amu). Calculate the average atomic weight. (39.1 amu)

3. The principal isotope of chlorine is 2.9140 times as massive as an atom of ^{12}C. What is its atomic mass? (34.968 amu)

SECTION 2: THE MOLE CONCEPT

The gram molecular weight (GMW) of a substance is the number of grams corresponding to its molecular weight. For example, the GMW of KCl is 74.552 g (K = 39.102 + Cl = 35.45). This quantity is often referred to as a mole of the substance.

A mole of any substance (either atomic, ionic, or covalent) contains the same number (6.023×10^{23}) of particles. The number is referred to as Avogadro's number.

Examples:

Substance	Gram Molecular Weight*	Number of Particles	Number of Moles
CCl_4	153.8	6.023×10^{23}	1
SO_3	80.06	6.023×10^{23}	1
H_2SO_4	98.07	6.023×10^{23}	1
Al	26.98	6.023×10^{23}	1
H_2	2.016	6.023×10^{23}	1

*Shown to 4 significant figures.

Laboratory

The student, using both triple beam balances and analytical (±0.0001 g) balances should weigh various elements and compounds in assigned "mole" quantities. For example, the actual mass of exactly 0.5 mole of aluminum, lead, H_2O, and NaCl could be visually compared. This also provides an excellent opportunity to train the student in proper weighing techniques.

Student Problems

1. What is the weight of a mole of:

   a) boric acid (H_3BO_3)? (61.83 g)

   b) potassium bromide (KBr)? (119.0 g)

2. How many molecules would be in three moles of carbon tetrachloride? (18.069×10^{23} molecules.)

3. How many moles are in 34.00 g of ammonia (NH_3)? (2 moles)

SECTION 3: THE PERIODIC TABLE

The periodic table is simply a display of the elements with respect to their chemical similarities. All elements in the same family (vertical columns) have identical outer-electron arrangements, thus similar chemical properties. For example, the alkali metal family (vertical column IA-elements: H, Li, Na, K, Rb, Cs, and Fr) contains all of the A-family elements which have single electrons in their outermost shells.

Periodic Table of the Elements

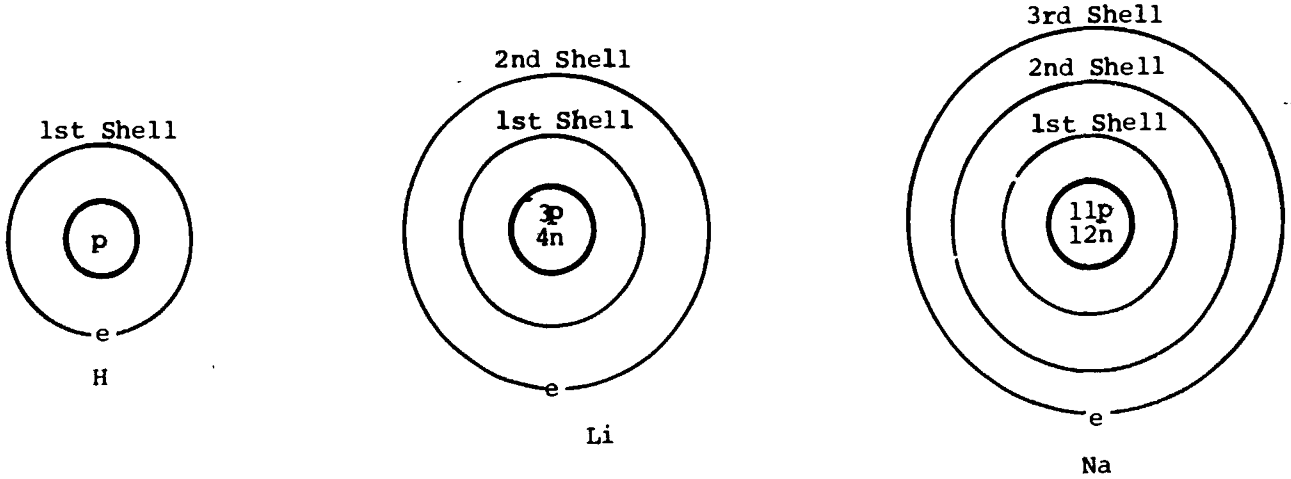
IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1.008 H 1																	4.003 He 2
6.94 Li 3	9.01 Be 4											10.61 B 5	12.01 C 6	14.01 N 7	16.00 O 8	19.00 F 9	20.18 Ne 10
22.99 Na 11	24.31 Mg 12											26.98 Al 13	28.09 Si 14	30.97 P 15	32.06 S 16	35.45 Cl 17	39.95 Ar 18
39.10 K 19	40.08 Ca 20	44.96 Sc 21	47.90 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.71 Ni 28	63.55 Cu 29	65.37 Zn 30	69.72 Ga 31	72.59 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36
85.47 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	98.91 Tc 43	101.1 Ru 44	102.9 Rh 45	106.4 Pd 46	107.9 Ag 47	112.4 Cd 48	114.8 In 49	118.7 Sn 50	121.8 Sb 51	127.6 Te 52	126.9 I 53	131.3 Xe 54
132.9 Cs 55	137.3 Ba 56	138.9 La* 57	178.5 Hf 72	180.9 Ta 73	183.9 W 74	186.2 Re 75	190.2 Os 76	192.2 Ir 77	195.1 Pt 78	197.0 Au 79	200.6 Hg 80	204.4 Tl 81	207.2 Pb 82	209.0 Bi 83	(210) Po 84	(210) At 85	(222) Rn 86
(223) Fr 87	226.0 Ra 88	(227) Ac† 89	(257) Ku 104	(260) Ha 105													

140.1 Ce 58	140.9 Pr 59	144.2 Nd 60	(145) Pm 61	150.4 Sm 62	152.0 Eu 63	157.3 Gd 64	158.9 Tb 65	162.5 Dy 66	164.9 Ho 67	[illegible] Er 68	[illegible] Tm 69	173.0 Yb 70	175.0 Lu 71
232.0 Th 90	231.0 Pa 91	238.0 U 92	237.0 Np 93	(242) Pu 94	(243) Am 95	(247) Cm 96	(249) Bk 97	(251) Cf 98	(254) Es 99	(253) Fm 100	(256) Md 101	(254) No 102	(257) Lr [illegible]

* Lanthanide Series
† Actinide Series

The atomic number of the element is simply its number of protons. For example, nitrogen which has an atomic number of 7 will always have 7 protons.

Isotopes: Two atoms within an element are said to be isotopes when they have the same number of protons but differ in their number of neutrons. Isotopes of an element have identical chemical properties but may differ in their degree of nuclear stability.



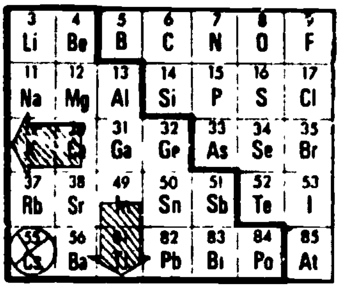
The periodic table consists of A-family and B-family elements. The A-family elements, also called the representative elements, obey the periodic law in that their chemical properties are predictable with other elements having similar outer electron arrangements. The B-family (transitional) elements all have similar chemical properties since their inner electron shells are filling. As an example, consider the elements in the period containing elements 21 to 30, scandium to zinc. Their 4th shells contain only two electrons while their 3rd shells are gaining from 1 to 10 electrons as the atomic number increases. In general, it can be stated that the B-family elements are those that have their next to outermost electron shells incompletely filled.

A metal can be defined as any element that tends to give up its outer electrons in a chemical reaction. Any element above helium whose atoms have one, two, or three electrons in their outermost shell can be considered a metal. Metals are good conductors of electricity, heat, etc.

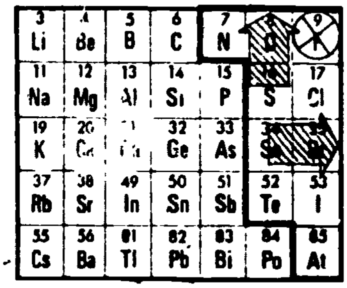
Nonmetals are elements that tend to accept additional electrons into their outermost shells in a chemical reaction. Any element whose atoms have five, six, seven, or eight electrons in their outside shells is considered a nonmetal. They are poor conductors of electricity and heat.

The following charts show the trends of the A-family (representative) elements to act as metals and/or nonmetals. The arrow shows the increase in the property with the (X) indicating the element that has the property to the greatest degree.

Tendency to act as a <u>metal</u>



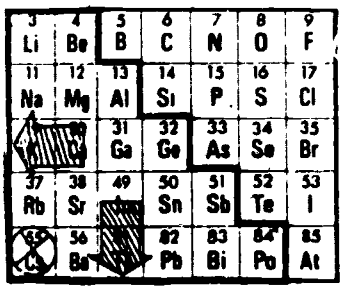
Tendency to act as a <u>nonmetal</u>



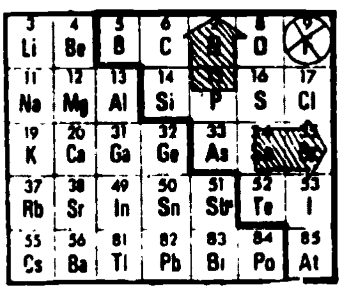
A <u>cation</u> can be defined as an atom or group of atoms with an over-all positive charge. Most metals, because of their tendency to lose electrons, form cations. For example, an aluminum atom becomes an Al^{3+} ion in a chemical reaction. An example of a polyatomic cation would be the ammonium (NH_4^+) ion.

An <u>anion</u> can be defined as an atom or group of atoms with an over-all negative charge. Most nonmetals, because of their tendency to gain electrons, form anions. Sulfur (a nonmetal) could become the S^{2-} ion in a chemical reaction. The carbonate (CO_3^{2-}) ion is an example of a polyatomic anion.

Tendency to become a <u>cation</u>



Tendency to become an <u>anion</u>



Laboratory

The student should perform "flame tests" on various salts to demonstrate the characteristic spectra from different atoms. This exercise could be extended by having students identify unknown salts by their characteristic visible light spectrum.

Student Problems

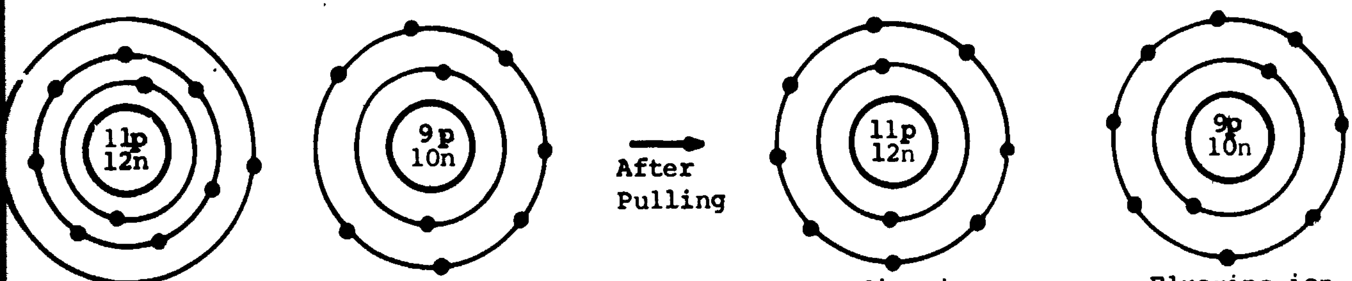
1. Predict what change will result on a barium atom upon being converted to a barium ion.
2. What do the elements Al, Ga, and In have in common?
3. How many electrons does an oxygen atom have in its valence shell?

Chapter III

Inorganic Chemistry

SECTION 1: ELECTRONEGATIVITY

When two elements differing in electronegativity (attraction for electrons) are brought together, the one with greater affinity pulls away electrons from the other element. This "pulling" of electrons creates a negative ion (from the more electron-attracting element) and a positive ion (from the less electron-attracting element).



dium atom
ectronegativity = 0.9

Fluorine atom
electronegativity = 4.0

Sodium ion
Na^+

Fluorine ion
F^-

Electronegativity of the representative elements (Groups 1A, 2A, 3A, 4A, 5A, 6A, and 7A) increases from left to right across a period and from bottom to top within a group.

Electronegativities for Representative Elements

Increasing Electronegativity →

		H 2.1					
Li 1.0	Be 1.5		B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2		Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0		Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0		In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9						

↑ Increasing Electronegativity

For example, sulfur (S) is more electronegative than magnesium (Mg) since it appears farther to the right on the table, even though it is in the same period.

Student Problems

1. Which element potassium (K) or aluminum (Al) would donate an electron the most readily to fluorine? (Potassium)

2. If carbon (C) and beryllium (Be) were to form an ionic compound,which element would be con-`dered to form the negative ion? (Carbon)

SECTION 2: ELECTRON-DOT STRUCTURES

"Lewis" or electron-dot structures are often used to depict what happens during ion or bond formation. The electron-dot picture for an element is simply the symbol for the element with dots drawn around it to represent the valence (outermost) shell of electrons. Elements from groups 1A, 2A, and 3A form positive ions by losing the number of electrons equal to their group numbers. Elements from groups 6A and 7A form negative ions by gaining the number of electrons equal to 8 minus the group number.

The electron-dot picture for the chlorine atom and a chlorine ion should be:

:C̈l· (with dots above and below)	:C̈l: (with dots above and below)
Chlorine Atom	Chlorine Ion

Chlorine is in group 7A and thus has 7 valence electrons. Since it is a nonmetal, it tends to gain electrons and when ionized has developed a -1 charge. The electron dots are normally shown in pairs but their positioning is not significant in these Lewis structures. The electron-dot pictures for the ions formed when barium and sulfur are brought together are shown below.

Ba: → Ba^{2+} (loses two electrons)

:S: → :S̈:$^{2-}$ (gains two electrons)

Thus:

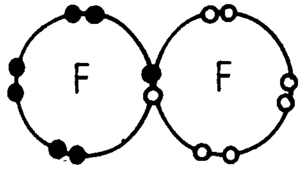
Ba :S̈: or simply BaS

Laboratory

The student should synthesize a pure compound of silver chloride starting with an exactly weighed mass of elemental silver. The silver metal is then dissolved in nitric acid and an excess of hydrochloric acid is added. The resulting precipitate should be used to calculate the empirical formula for silver chloride. (Modern Chemical Technology experiment 8-1 is a typical type of formula determination experiment.)

Student Problems

1. Draw the electron dot structure for fluorine gas (F_2).



2. How many valence electrons does the element carbon have?

SECTION 3: NAMING INORGANIC COMPOUNDS

It has been estimated that there are over 10,000 inorganic compounds known. It would be beneficial for the name of the compound to indicate something about its chemical structure. Thus, a systematic nomenclature (naming) has been developed.

The simplest inorganic compounds are those that contain only two elements called binary compounds. Most compounds can be named by following these rules:

1. Write the name of the cation or less electronegative element first. This will normally be a metal, except in those cases in which a polyatomic positive ion (for example NH_4^+) is involved. Write the name of the anion or more electronegative element last. The name will be formed in most binary compounds by dropping the usual ending of the elements name and adding "ide".

   A sodium and fluorine ionic compound would be called: sodium fluoride (NaF).

   Sulfur and calcium would be called calcium sulfide (CaS) while lithium and iodine would be called lithium iodide (LiI)

2. When a metal forms more than one ion, place the Roman numeral corresponding to the positive charge on the metal ion in parentheses after the name of the metal. Iron (Fe) is known to form both the Fe^{2+} and Fe^{3+} ions, thus the name iron chloride could be either $FeCl_2$ or $FeCl_3$. There would be no confusion with iron (II) chloride and iron (III) chloride respectively. An old naming system does exist in which the lower number metal ion is ended with "ous" and the higher ion as "ic". Thus ferrous (Fe^{2+}) and ferric (Fe^{3+}) names will be encountered.

3. When two elements involved may form more than one compound with each other, use Greek prefixes--di-, tri-, tetra-, penta-, hexa-, and so forth--to indicate how many atoms of each are in the molecule. Nitrogen and oxygen are two elements that combine in different proportions. For example, N_2O, NO_2 and N_2O_4 are all known to exist. Nitrogen oxide would not be descriptive enough but dinitrogen oxide, nitrogen dioxide, and dinitrogen tetraoxide would show the proportions.

4. When a polyatomic ion is involved, its name will be written as shown in the following table.

POLYATOMIC IONS

Ion	Name	Ion	Name
$C_2H_3O_2^-$	acetate	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	HCO_3^-	hydrogen carbonate
CO_3^-	carbonate	OH^-	hydroxide
ClO_3^-	chlorate	NO_3^-	nitrate
Cl^-	chloride	NO_2^-	nitrite
ClO_2^-	chlorite	MnO_4^-	permanganate
ClO^-	hypochlorite	PO_4^{3-}	phosphate
ClO_4^-	perchlorate	SO_4^{2-}	sulfate
CrO_4^{2-}	chromate	SO_3^{2-}	sulfite
CN^-	cyanide		

Naming Inorganic Acids

An important group of inorganic compounds having the general formula H_aY dissolve in water to form H+ ions and Y^{a-} ions. These are called acids and are not customarily named according to the rules used for other compounds.

For ternary acids, (three elements involved) name the Y portion by writing the name of the Y^{a-} ion (see polyatomic ions table), making the following changes in endings from ion name to acid name: change "-ate" to "-ic"; change "-ite" to "-ous"; then write the word "acid". For example, a solution containing HNO_3 would not be called hydrogen nitrate. It would be called nitric acid. The "ate" ending of the nitrate (Y^{a-}) ion would be dropped and the "ic" ending added.

For binary acids (two elements involved), write the prefix "hydro-" followed by the name of the Y portion, changing the usual ending to "-ic"; then write the word "acid". An example would be aqueous (dissolved in water) HCl which is properly called hydrochloric acid. This compound when pure (as opposed to dissolved in water) could be called "hydrogen chloride".

The Special Nomenclature of Acids

Formula	Standard Nomenclature	Acid Nomenclature
HCl	hydrogen chlorIDE	HYDROchlorIC acid
HCN	hydrogen cyanIDE	HYDROcyanIC acid
HBr	hydrogen bromIDE	HYDRObromIC acid
HF	hydrogen fluorIDE	HYDROfluorIC acid

When the suffix is IDE, the acid is HYDRO-root-IC acid.

HNO_2	hydrogen nitrITE	nitrOUS acid
H_2SO_3	hydrogen sulfITE	sulfurOUS acid
H_3PO_3	hydrogen phosphITE	phosphorOUS acid
H_3AsO_3	hydrogen arsenITE	arsenOUS acid

When the suffix is ITE, the acid is root-OUS acid.

HNO_3	hydrogen nitrATE	nitrIC acid
H_2CrO_4	hydrogen chromATE	chromIC acid
H_2CO_3	hydrogen carbonATE	carbonIC acid
H_2SO_4	hydrogen sulfATE	sulfurIC acid

When the suffix is ATE, the acid is root-IC acid.

$HClO$	hydrogen HYPOchlorITE	HYPOchlorOUS acid
$HBrO$	hydrogen HYPObromITE	HYPObromOUS acid
$HMnO_4$	hydrogen PERmanganATE	PERmanganIC acid
$HClO_4$	hydrogen PERchlorATE	PERchlorIC acid

The prefixes HYPO and PER remain in the acid nomenclature.

A few examples of applying these inorganic acid nomenclature rules would be:

H_2SO_3 -- sulfurous acid

H_2SO_4 -- sulfuric acid

HCl -- hydrochloric acid

HBr -- hydrobromic acid

H_2CrO_4 -- chromic acid

HNO_2 -- nitrous acid

$HClO_4$ -- perchloric acid

Laboratory

The student should be able to write the formula of most common inorganic compounds by simply reading the name. An excellent exercise would be to have each student inventory at least 200 inorganic compounds found in the stockroom and list both the proper name and formula.

Student Problems

1. Name the compounds:

   a) $LiBr$

   b) H_2SO_3 (in water)

   c) $Ca(ClO)_2$

   d) NaCn

   e) NH_4S

Chapter IV

Nuclear Chemistry

SECTION 1: FISSION AND FUSION

Radioactivity is a process in which energy is released by the nucleus of an atom as it achieves greater stability. Nuclear radiation is most often accomplishe by a loss of some nuclear material, but rarely involves the electrons of an atom.

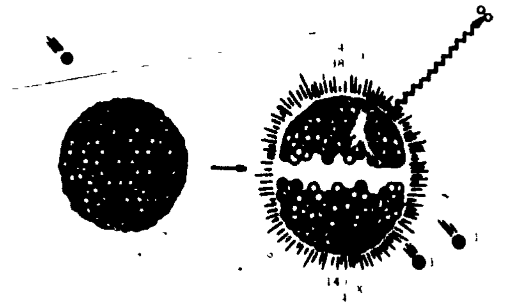
Elements that spontaneously emit energy without first absorbing energy are said to be naturally radioactive. There are two main nuclear reaction categories: fusion and fission. The process whereby light nuclei combine to form heavier nuclei is called fusion. Fusion reactions are responsible for the energy released by our sun and other stars. Because the fusing particles are positively charged, high kinetic energies (perhaps equivalent to 10,000,000° C) are required to bring them sufficiently close to interact.

$$^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n + \text{energy}$$

Notice that two isotopes of hydrogen have "fused" or combined to produce the element helium and energy. The superscript represents the atomic mass and the subscript indicates the atomic number in all nuclear equations.

In fission, a heavy nucleus breaks up into two smaller nuclei and several neutrons.

FISSION



$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{236}_{92}U \rightarrow ^{140}_{54}Xe + ^{94}_{38}Sr + 2\ ^{1}_{0}n$$

Most fission reactions are initiated by neutrons and are extremely fast. More than one neutron is often produced by fission. Thus, an explosive chain reaction is possible. Control is usually achieved by limiting the number of neutrons through the use of neutron absorbing materials. Many safety problems are associated with nuclear reactors, including venting of radioactive gases, the possibility of a major explosion with the subsequent emission of radioactive materials into the environment, and the storage of highly radioactive waste material.

SECTION 2: TYPES OF RADIATION

Alpha radiation is composed of a stream of alpha particles from the nuclei. Alpha particle emission is characteristic of large nuclei. The alpha particle is identical to a doubly-ionized helium ion, being composed of two protons and two neutrons. A loss of an alpha particle causes a transformation of the elemental atom as well as a loss in mass.

The alpha particle is represented as ${}^{4}_{2}He$. Examples of alpha particle emission are:

$${}^{234}_{90}Th \rightarrow {}^{230}_{88}Ra + {}^{4}_{2}He$$

$${}^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$$

Notice that the sum of the masses and charges on both sides of the equation are equal.

Alpha radiation is not as hazardous as the others to be discussed. The massive alpha particle is relatively slow moving and its range in air is only a few centimeters. However, an alpha-emitting substance that is breathed in or ingested is damaging because it destroys living cells by ionizing atoms of protein molecules which compose the cell.

Beta radiation consists of a stream of electrons (if negative) or a stream of positrons (if positive).

The beta particle is produced when a neutron decays in the nucleus into a proton and a beta.

The symbol for positive and negative beta particles are ${}^{0}_{-1}e$ for the electron and ${}^{0}_{+1}e$ for the positron. The superscript of zero indicates the mass of a beta particle. The electron produced by nuclear decay is identical to the normal electrons outside the nucleus. An example of beta emission is:

$${}^{234}_{90}Th \rightarrow {}^{234}_{91}Pa + {}^{0}_{-1}e$$

Note once again that the sums of the masses and charges are equal. Because of the negligible mass of the particle, it has only the slightest ionizing power. The range of beta particles depends on their energy, but the average range is limited to a few meters. Beta radiation as a hazard is not great except for the possible exposure to a direct beam, in which case the ionization of protein atoms could be serious.

Gamma radiation is produced by nuclei in a highly excited state. Gamma radiation is electromagnetic. The wavelength is extremely short (a small fraction of a meter) and the photon of this radiation has a very high energy. Its range may be measured in light years as it streaks through the void of the universe from distant stars. The gamma radiation in this case is commonly called cosmic radiation. The penetrating power of these cosmic rays, without mass or change, is remarkable. Neither lead walls nor deep caves below the earth can fully block the effect of cosmic radiation. In many cases, when a nucleus is left in an excited state as a result of change, gamma rays are emitted as the nucleus proceeds toward the ground state. The usual gamma rays of the nuclear decay of familiar isotopes are not as penetrating as cosmic radiation. Gamma radiation is very hazardous. The photons have sufficient energy to destroy the molecules of living protoplasm by gross alteration or "burning."

A SUMMARY OF RADIATION TYPES

Unit	Common Symbol	Modern Symbol	Mass (amu)	Charge
alpha	α	${}^{4}_{2}He$	4	+2
beta	$\beta-$	${}^{0}_{-1}e$	0	-1
positron	$\beta+$	${}^{0}_{+1}e$	0	+1
gamma	γ	γ	0	0

SECTION 3: RATE OF RADIOACTIVE DECAY

Radioactive nuclei have different stabilities and they disintegrate at different rates. Each radioactive process is characterized by a half-life. The half-life is the time required for one-half of a sample of radioactive atoms to decay. During the interval of one half-life, one-half of the number of radioactive atoms originally present decay, and one-half of the original sample remains unchanged. Suppose one starts with a 1.000 g sample of tritium, a hydrogen isotope that decays by beta emission with a half-life of 12.3 years. After 12.3 years, only 0.500 g would remain. After 24.6 years, only 0.250 g of the original sample would remain. After 36.9 years, only 0.125 g would remain, and so on.

The half-life values for various common isotopes is given in the table below. A column of relative isotopic abundance is provided to indicate the % composition of these elements. For example, pure naturally-occurring uranium (U) contains 99.28% ^{238}U and 0.72% ^{235}U.

NATURALLY OCCURRING RADIOACTIVE ISOTOPES

Isotope	Type of Disintegration	Half-life (yr)	Relative Isotopic Abundance (%)
^{3}H	β	12.3	0.00013
^{87}Rb	β	6.2×10^{10}	27.8
^{115}In	β	6×10^{14}	95.8
^{144}Nd	α	$\sim 5 \times 10^{15}$	23.9
^{147}Sm	α	1.3×10^{11}	15.1
^{176}Lu	β	4.6×10^{10}	2.60
^{190}Pt	α	$\sim 1 \times 10^{12}$	0.012
^{238}U	α	4.5×10^{9}	99.28
^{232}Th	α	1.4×10^{10}	100
^{14}C	β	5730	trace
^{235}U	α	7.13×10^{8}	0.72

Human Exposure and Safety

The effects of high energy radiation may be measured in a variety of ways, using several kinds of units of radiation such as roentgens, REM (roentgen equivalent

in man), and rads. A roentgen (r) is defined as the amount of radiation that will produce one electrostatic unit of ions per cubic centimeter volume. The milliroentgen, abbreviated, as mr, is 1/1000 of a roentgen. A rad is equivalent to one erg of energy absorbed per gram of tissue. Roentgens and rads are usually chosen as a measure of radioactive dosage; some effects are described in the following charts.

Common Radiation Exposures in Roentgens (r)

Man Made	
Source	*Exposure*
Fluoroscope	5 to 400 r
Chest X-ray	10 mr to 1 r
Dental X-ray	20 to 65 r
Pregnancy X-ray	20 to 65 r
Barium studies (X-ray)	10 to 20 r per min
Diagnostic studies for Heart disease for Acne treatment for Malignant tumor (local dose)	 140 r 500 to 1,000 r per treatment 3,000 to 7,000 r
Wrist watch	40 mr per year
Luminous dials in airplanes	1.3 r per year
Uranium mine	5.6 r per year
Fallout	100 mr per year
Television sets (3 ft away	0.04 mr per hour

Natural	
Source	*Exposure*
Cosmic rays (sea level)	40 mr per year
Granite rock	100 mr per year
Ocean water	50 mr per year
Average soil	30 to 80 mr per year
Radon in air	120 mr per year
$^{40}_{19}K$ in the body	20 mr per year
$^{226}_{88}Ra$ (bone)	40 mr per year
$^{40}_{19}K$ from people in packed crowds	2 mr per year
Low grade uranium ore	2.8 r per year
$^{14}_{6}C$ in the body	1 mr per year

Health Effects of Radiation Exposure

Dose in Rads	Probable Effect on Man
0- 50	Possible genetic mutation
50-100	Headache, dizziness, listlessness
100-200	Radiation sickness and hair loss
200-500	Severe bleeding and tissue destruction
Over 500*	Death

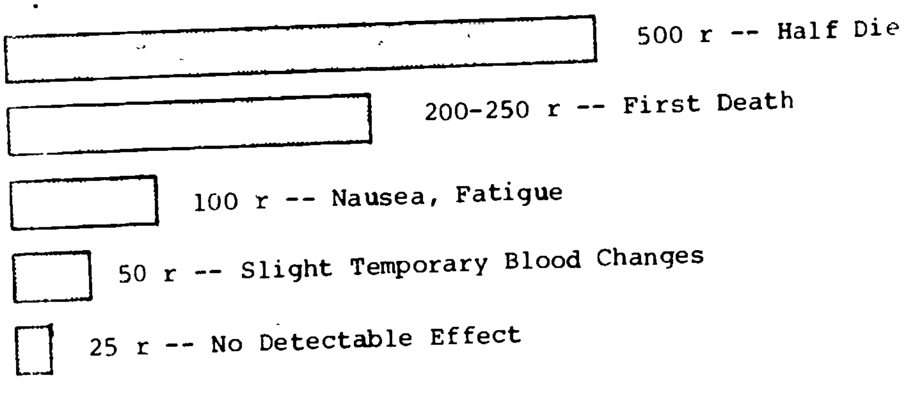
* This is known as a lethal dose (LD).

Whole Body Radiation vs. Isolated Exposure

It is impossible to say how many roentgens it will take to kill any specific individual because we all vary in our resistance to any attack upon the body, whether it is by radiation, electricity, poison, injury, disease, etc. It is quite certain, however, that no human being could survive 1,000 r of total body radiation delivered in a short space of time. Both the total dose, the concentration and the time are important considerations. The effect of 1,000 r of radiation delivered to the total body is by no means the same thing as 1,000 r delivered to a small portion of the body any more than a first degree burn of the palm of the hand is the same thing as a third degree burn of a large area of the body.

Generally, 24 hours or less is considered a short time. The ability of the body to withstand any injury is increased if the same amount of injury given to the body is spread over a longer period of time. Whiskey can be poison, but many people can drink an ounce of whiskey each evening before dinner over an extended period of time without apparent harm. If a person drinks a fifth of whiskey at one time, he could die of alcoholic poisoning because the body has not been given sufficient time to recover from the effect of the alcohol.

The radiation dose it takes to kill one specific individual is not a good measure of the fatal dose to others because of individual differences. The term used is the median lethal dose, or LD/50. This is the dose required to kill 50% of the subjects. The LD/50 for penetrating external radiation is about 500 r delivered to the total body in 24 hours or less. This means that if a representative sample of the population were subjected to 500 r of total body radiation within a 24-hour period, approximately 50 percent of these people would die, and the other 50 percent would recover. The effects of less radiation are shown in the following graph.



Effects of External Radiation for Total Body Exposure within a 24-hour Period

As radiation penetrates the body, there is no immediate associated sensation. The shorter the period of exposure, the less likely the radiation will affect tissue. When it is essential to be exposed to radiation, the best method of protection is distance. The intensity of the radiation falls off as the inverse square of the distance from the source. This statement is known as the inverse-square law and is expressed arithmetically as follows:

$$I = \frac{I_0}{r^2}$$

where I_0 is the original intensity of a source of radiation and I is its intensity at a distance r. Thus, if a radiation source registers 1,000 rads on a Geiger counter held 1 ft from the source, it will register 250 rads when held 2 ft from the source:

$$I = \frac{1{,}000 \text{ rads}}{4}$$

$$= 250 \text{ rads}$$

Student Problems

1. Write the equation for the fission of a uranium-235 atom into molybdenum-102 and tin-131 nuclei.

2. Describe the difference between fission and fusion.

3. Predict the resulting element if a $^{226}_{88}Ra$ atom emitted an alpha particle.

   ($^{222}_{86}Rn$)

4. Which type of radiation is the most hazardous to human beings?

5. Given that the half-life of lead-214 is 26.8 months, calculate the quantity of radioactive lead-214 that would remain if 10 grams stood for 80.4 months.

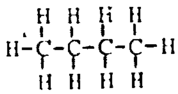
   (1¼ grams)

Chapter V

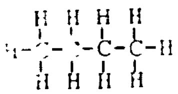
Organic Chemistry

SECTION 1: ALIPHATIC HYDROCARBONS

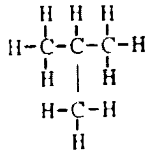
There are more organic (carbon-containing) compounds known than any other type. The simplest organic compounds contain only "hydrogen" (H) and "carbon" (C) atoms and are thus called hydrocarbons. The simplest hydrocarbons, the alkanes, have the general formula C_nH_{2n+2} where n represents the number of carbon atoms and 2n+2 represents the number of attached hydrogen atoms. For example, butane (C_4H_{10}) contains 4 carbon atoms; thus it must contain 10 hydrogen atoms (2 x 4 + 2 = 10). It should be pointed out that carbon atoms always require 4 bonds and hydrogen atoms are capable of making only one bond. The structural formula for butane is this:



The naming of hydrocarbons would be very easy if things were all this simple. However, C_4H_{10} can represent two possible geometric structures or isomers (compounds with the same general formula but different structures):



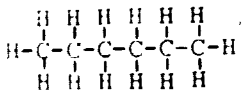
and

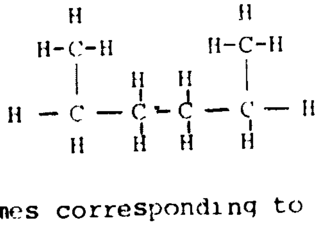


The number of possible isomers that a formula may represent increases rapidly with increasing numbers of carbon and hydrogen atoms. C_8H_{18} can represent 18 different isomers.

The basic rules for naming hydrocarbons (according to the International Union of Pure and Applied Chemistry - IUPAC) are as follows:

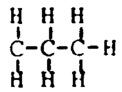
1. Name the hydrocarbon by selecting the name of the longest continuous chain of carbons found in the structure. For example, both of the following compounds contain a continuous chain of 6 carbon atoms with no branches, and both are hexane.



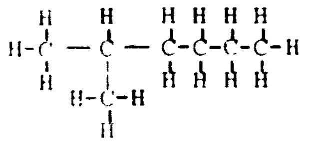


See the table for the hydrocarbon names corresponding to different numbers of carbon atoms.

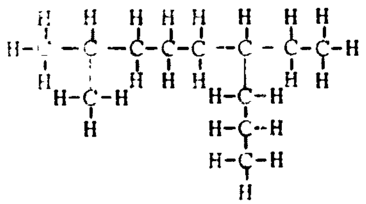
2. All carbon-hydrogen groups not counted in the continuous chain should be named as alkyl groups. An alkyl group is simply a hydrocarbon with one hydrogen atom removed. Its name is derived by dropping the "ane" ending on the parent hydrocarbon and adding "yl". For example, the alkyl group



derived from propane would be called propyl. The alkyl group names should preceed the name of the longest continuous chain.

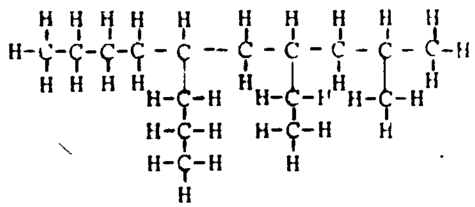


methyl-hexane



methyl-ethyl-nonane

3. The carbon atoms in the longest continuous chain are numbered in an order which will locate the attached alkyl groups using the smallest numbers possible. For example,



2-methyl-4-ethyl-6-propyldecane

ALKANE TABLE

Number of Carbons	Name of Hydrocarbon	Hydrocarbon Formula
1	methane	CH_4
2	ethane	C_2H_6
3	propane	C_3H_8
4	butane	C_4H_{10}
5	pentane	C_5H_{12}
6	hexane	C_6H_{14}
7	heptane	C_7H_{16}
8	octane	C_8H_{18}
9	nonane	C_9H_{20}
10	decane	$C_{10}H_{22}$

Alkenes are hydrocarbons which contain one or more double bonds. They are said to be unsaturated hydrocarbons inasmuch as they contain less than the maximum number of hydrogens that could be accommodated by the carbons present. The general formulas for the alkenes and cycloalkenes (cyclic structure) are:

open chain alkenes C_nH_{2n}

cycloalkenes C_nH_{2n-2}

Nomenclature for the simple alkenes is also based on the names given in the alkane table, but the suffix -ene rather than -ane is used. Alkene nomenclature also requires that when there may be confusion otherwise, a numerical prefix be used to show the location of the double bond along the carbon chain. The carbons are numbered from the end of the chain that gives the lowest number to a carbon that is doubly bonded, and the number of only one of the pair of carbons involved in the bond is needed.

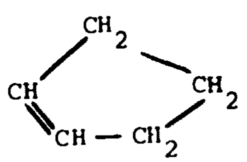
Examples

$CH_2{=}CHCH_2CH_2CH_3$

1-pentene

$CH_3CH{=}CHCH_2CH_3$

2-pentene



cyclopentene

Alkynes are hydrocarbons containing one or more triple bonds. Alkynes are even more highly unsaturated than the alkenes, as can be seen from their formulas. The general formulas for alkynes and cycloalkynes are:

open chain alkynes	C_nH_{2n-2}
cycloalkynes	C_nH_{2n-4}

The naming of alkynes follows the same approach taken for alkenes. The -ane endings (see alkane table) are replaced with -yne endings, and numerical prefixes are used when needed to locate the position of the triple bond.

Examples

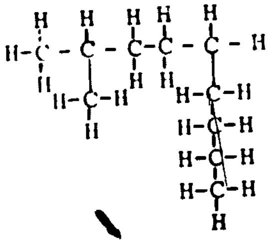
$CH CCH_2CH_2CH_3$	$CH_3C^{-}CCH_2CH_3$
1-pentyne	2-pentyne

Student Problems

1. If a butyl and a hexyl alkyl group were combined, an isomer of which hydrocarbon would be created?

   a) nonane
   b) butylhexane
   c) decane
   d) hexylbutane
   e) hexane

2. Using the IUPAC system, select the name of the following hydrocarbon.



   a) 2-methylnonane
   b) decane
   c) 2-methyl-5-butylpentane
   d) 1-butyl-4-methylpentane
   e) 8-methylnonane

3. Which of the following hydrocarbons would have the greatest number of isomers?

   a) ethane
   b) propane
   c) butane
   d) pentane
   e) hexane

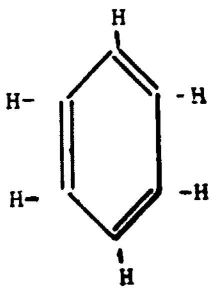
4. Name this compound.

```
  | |     | | |
 -C-C-C≡C-C-C-C-
  | |     | | |
```

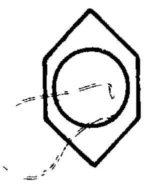
5. The compound shown above is an alkane, alkene, or alkyne?

SECTION 2: AROMATIC HYDROCARBONS

Aromatic compounds are special, cyclic, unsaturated hydrocarbons containing the benzene ring. Benzene (C_6H_6) can be depicted several ways but the abbreviated formula is most common.



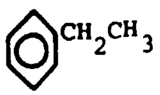
Structural Formula



Abbreviated Formula

Benzene, like other hydrocarbons, may have other atoms (or groups) substituted for a hydrogen on the C-skeleton. When a single alkyl group is substituted for hydrogen, we name the resulting compound by placing the name of the alkyl group [illegible] the word "benzene."

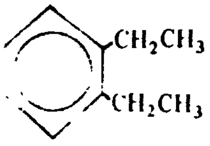
Example



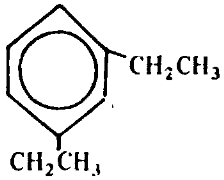
ethylbenzene

When two groups are substituted on the ring, the possible isomers are commonly distinguished from one another by using the prefixes ortho-, meta-, or para- (or simply o-, m-, p-).

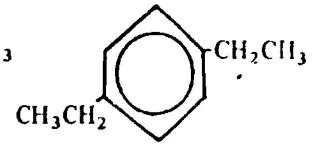
Examples



o-diethylbenzene (adjacent)



m-diethylbenzene (alternate)



p-diethylbenzene (diagonal)

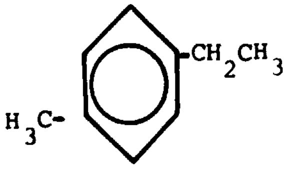
When more than two groups are substituted on the ring, the ring positions are numbered (giving the lowest possible numbers to carbons with substituents) and these numbers used as prefixes to distinguish among isomers. This numbering approach may also be used when only two substituents appear on the ring. For example, the aromatic compounds shown above would be named 1,2-diethylbenzene; 1,3-diethylbenzene; 1,4-diethylbenzene, respectively.

Laboratory

The students should construct using molecular kits, different alkanes, alkenes, alkynes, cyclic and aromatic hydrocarbons. This laboratory exercise could involve the entire class with the construction of various isomers, for example, the building of all 18 isomers of octane.

Student Problems

1. Name this compound using both the ortho, meta, and para system and the IUPAC (numerical) system.



2. Draw the compound: 1-ethyl 2,3-dipropylbenzene.

SECTION 3: FUNCTIONAL GROUPS

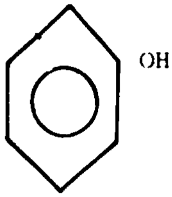
We may think of a typical organic compound as consisting of a carbon skeleton covered with a relatively inert hydrogen "skin" and possessing localized sites of reactivity - the functional groups that give organic compounds their chemical personalities.

The alkanes consist solely of C and H atoms, which differ little in electronegativity and are not very reactive. However, a hydrocarbon can be made more reactive by replacing hydrogen atoms with one or more of the following functional groups. The R is used to represent any alkane or aromatic compound.

SOME TYPES OF SIMPLE ORGANIC COMPOUNDS

Functional Group	General Formula of Compound	Type of Compound
-X (F,Cl,Br,I)	R-X	halide
-OH	R-OH	alcohol
-O-	R-O-R	ether
$-NH_2$	$R-NH_2$	amine
$-\overset{\overset{O}{\Vert}}{C}-H$	$R-\overset{\overset{O}{\Vert}}{C}-H$	aldehyde
$-\overset{\overset{O}{\Vert}}{C}-$	$R-\overset{\overset{O}{\Vert}}{C}-R$	ketone
$-\overset{\overset{O}{\Vert}}{C}-OH$	$R-\overset{\overset{O}{\Vert}}{C}-OH$	carboxylic acid
$-\overset{\overset{O}{\Vert}}{C}-O-$	$R-\overset{\overset{O}{\Vert}}{C}-O-R$	ester
$-\overset{\overset{O}{\Vert}}{C}-NH_2$	$R-\overset{\overset{O}{\Vert}}{C}-NH_2$	amide

For example R-OH could represent

-C-OH, -C-C-C-C-C-OH, or 

Using the common nomenclature system, many compounds can be named by simply stating the alkyl (hydrocarbon derivative) or aryl (aromatic derivative) group and the functional group or family. For example:

-C-C-C-C-Cl	butyl chloride
-C-C-OH	ethyl alcohol
-C-O-C-	dimethyl ether
-C-C-CH (with =O on the last C)	propyl aldehyde

The IUPAC system is much more comprehensive. A few general nomenclature rules might be helpful:

1. Alcohols can be named as substituted hydrocarbon derivatives by selecting the longest continuous hydrocarbon chain and changing the ending form "e" to "ol".

   Example: -C-C-C-C-C-OH

   pentanol

2. Aldehydes can be named by selecting the longest continuous hydrocarbon chain and changing the "e" ending to "al".

   Example:

   O
   ‖
   -C-C-C-

   propanal

3. Ketones are similar to aldehydes except that the -C- (carbonyl) group (with =O) appears within the hydrocarbon chain. The hydrocarbon's "e" ending is changed to "one".

   Example:

   O
   ‖
   -C-C-C-C-C-

   Pentanone

4. Carboxylic acid can be named by dropping the hydrocarbon "ane" ending and substituting "oic acid".

Example:

```
     O
  |  ‖
 -C-C-OH        ethanoic acid
  |
```

Laboratory

The student should be able to chemically identify and distinguish various organic unknowns by using "spot tests".

Student Problems

1. What type of organic compound would this structure represent?

```
 |     |  |
-C-O-C-C-
 |     |  |
```

2. Name this compound by both the common and IUPAC system.

```
 |  |  |  |
-C-C-C-C-OH
 |  |  |  |
```

3. Calculate the molecular weight for both pentanone and hexanone. (86, 100)

4. Draw the structure for octanal.

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